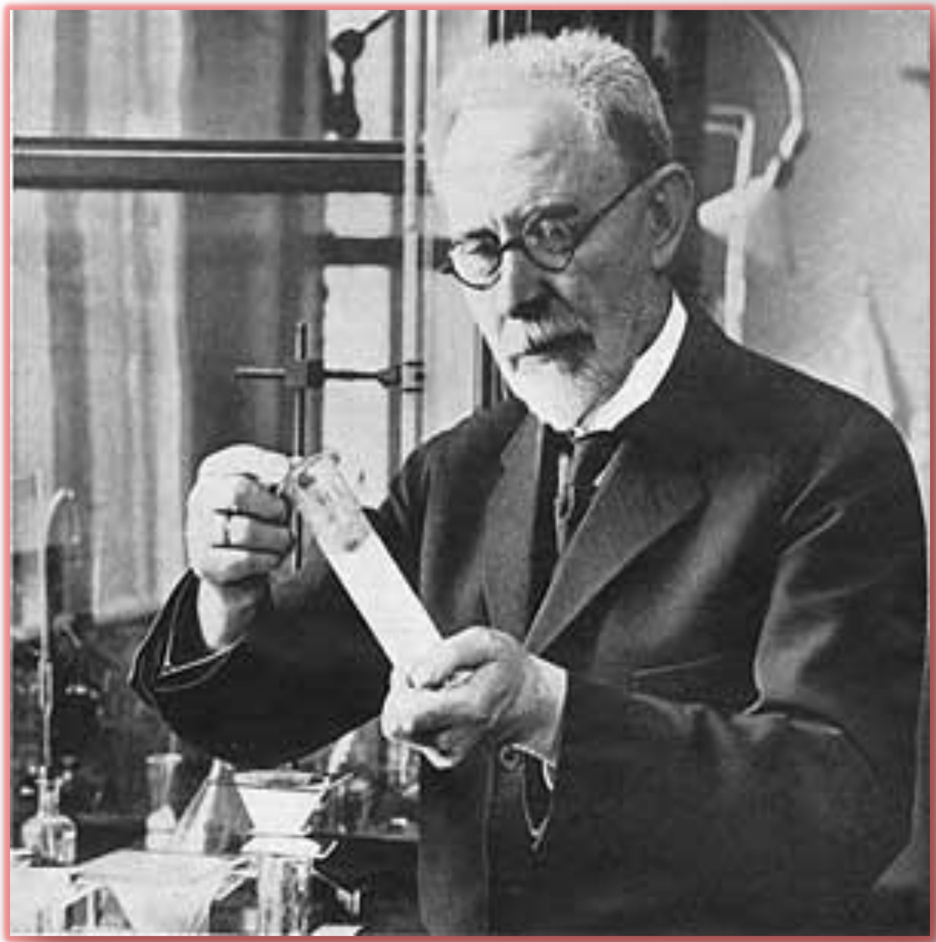
T08D03 – SL/HL Chem 2
Acids and Bases

The pH of Strong and Weak Acids and Bases

$[H^+]$	$[OH^-]$	pH	pOH
10^{-14}	1	14	0
10^{-13}	10^{-1}	13	1
10^{-12}	10^{-2}	12	2
10^{-11}	10^{-3}	11	3
10^{-10}	10^{-4}	10	4
10^{-9}	10^{-5}	9	5
10^{-8}	10^{-6}	8	6
10^{-7}	10^{-7}	7	7
10^{-6}	10^{-8}	6	8
10^{-5}	10^{-9}	5	9
10^{-4}	10^{-10}	4	10
10^{-3}	10^{-11}	3	11
10^{-2}	10^{-12}	2	12
10^{-1}	10^{-13}	1	13
1	10^{-14}	0	14

Ok, one more old dude!

- Søren Sørenson
  - Denmark
  - Biochemist
  - Early 20th century
  - Proposed the pH scale



Concentrations of $[H^+]$

- Actual concentrations of the hydronium ion, $[H^+]$, are often very small
- Therefore, Sørenson proposed a manipulation of the concentration of H^+ in a way that made the data much more simple to relate
- The pH scale is based on the logarithm of the concentration of H^+

Calculating pH

- The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration (in mol/L).
  - $pH = -\log[H_3O^+]$
  - $pH = -\log[H^+]$
- Therefore, the pH range of solutions are as follows:
  - Acidic Solutions, $pH < 7.0$
  - Basic Solutions, $pH > 7.0$
  - Neutral Solution, $pH = 7.0$

The pH scale is logarithmic

- Since pH 7 is neutral
  - pH 5 is 10 x more acidic than
    - pH 6
  - pH 4 is 100 x more acidic than
    - pH 6
  - pH 3 is 1000 x more acidic than
    - pH 6
  - pH 9 is 10 x more alkaline than
    - pH 8
  - pH 10 is 100 x more alkaline than
    - pH 8

Concentration of hydrogen ions compared to distilled water		Examples of solutions at this pH
10,000,000	pH = 0	battery acid, strong hydrofluoric acid
1,000,000	pH = 1	hydrochloric acid secreted by stomach lining
100,000	pH = 2	lemon juice, gastric acid, vinegar
10,000	pH = 3	grapefruit, orange juice, soda
1,000	pH = 4	tomato juice, acid rain
100	pH = 5	soft drinking water, black coffee
10	pH = 6	urine, saliva
1	pH = 7	"pure" water
1/10	pH = 8	sea water
1/100	pH = 9	baking soda
1/1,000	pH = 10	Great Salt Lake, milk of magnesia
1/10,000	pH = 11	ammonia solution
1/100,000	pH = 12	soapy water
1/1,000,000	pH = 13	bleaches, oven cleaner
1/10,000,000	pH = 14	liquid drain cleaner

The scale is courtesy of The Pacific Institute for the Mathematical Sciences

What is pOH?

- pOH is the opposite of pH.
- If pH goes down, pOH goes up.
- The pOH scale is:
  - Basic Solutions: pOH < 7.0
  - Acidic Solutions: pOH > 7.0
  - Neutral Solutions: pOH = 7.0
- pH + pOH = 14
- pOH = -log $[OH^-]$

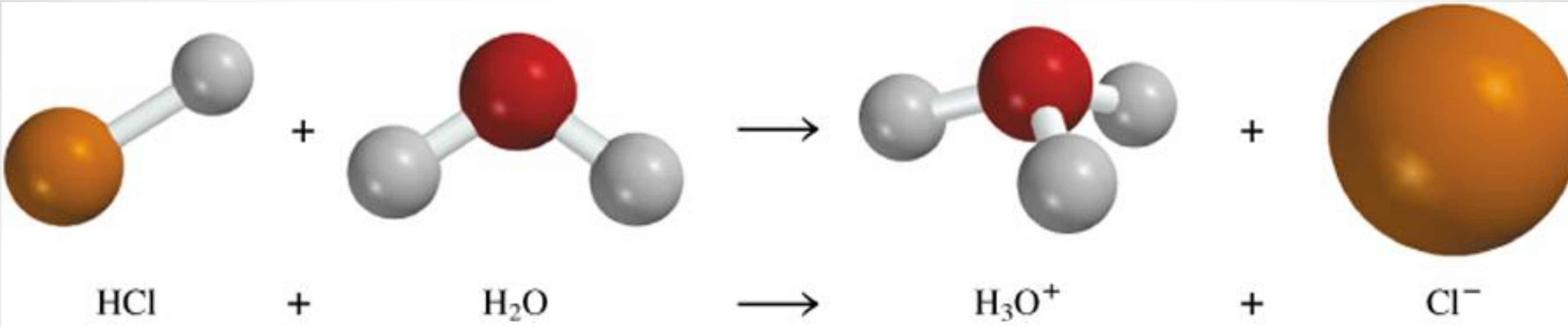
Find the pH of a 12M solution of HCl

- pH = -log (12M) = -1.07
- "Why is the pH of 12 Molar HCl 0 and not -1.07?"
- Any solution with a pH or pOH calculation that results in a negative number will have a pH or pOH of zero (0).

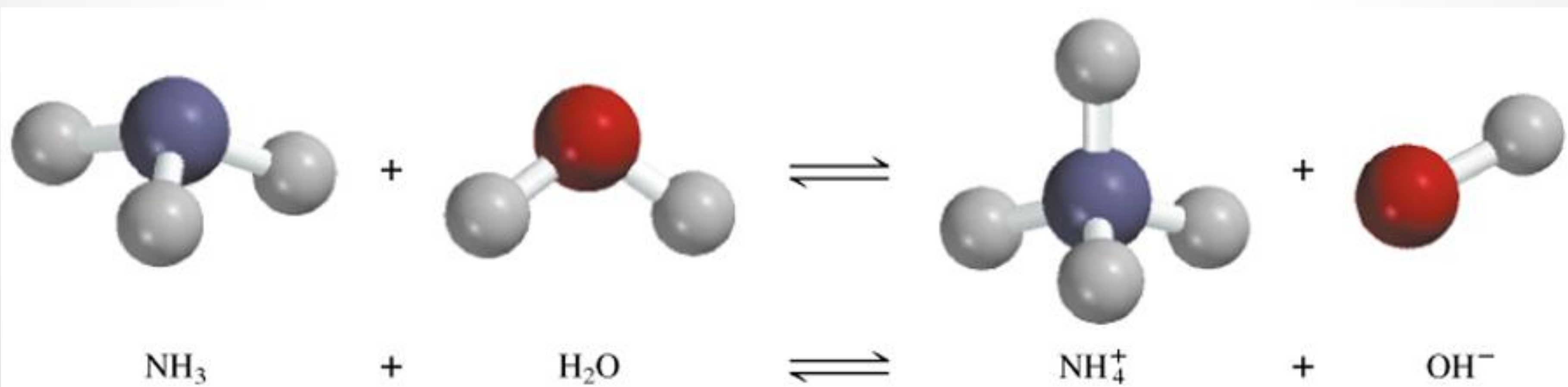
Aqueous Solutions of A&B

- What happens when you put an acid or a base into water?
- Each have the property of being electrolytes so will therefore dissociate
- Water itself can act as an acid or a base
  - $H^+ + H_2O \rightarrow H_3O^+$ (or $H^+ + OH^- \rightarrow H_2O$)
  - $H_3O^+ \rightarrow H^+ + H_2O$ (or $H_2O \rightarrow H^+ + OH^-$)

Arrhenius acid is a substance that produces H^+ (H_3O^+) in water

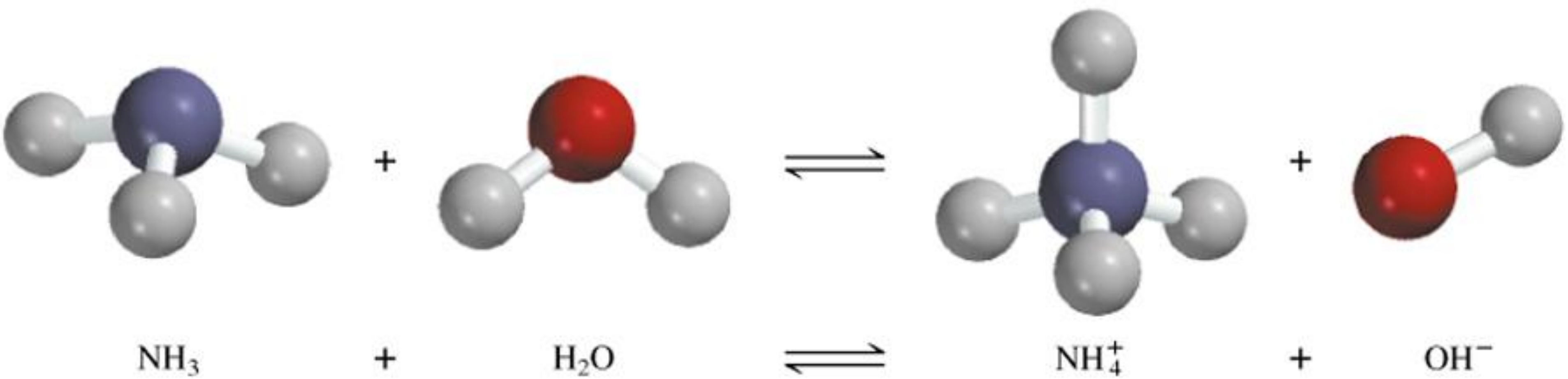


Arrhenius base is a substance that produces OH^- in water



A Brønsted acid is a proton donor

A Brønsted base is a proton acceptor



base acid conj. acid conj. base

Conjugate Acids and Bases

- Conjugate pairs are two substances that differ by one H^+ (they gain or lose one PROTON)

  HCl (acid) → Cl^- (conjugate base)

  NH_3 (base) → NH_4^+ (conjugate acid)

- When an acid loses a proton it becomes its conjugate base

  $H_2O \rightarrow OH^-$ or $HCl \rightarrow H^+ + Cl^-$

- When a base gains a proton it becomes its conjugate acid

  $H_2O \rightarrow H_3O^+$ or $NH_3 + H^+ \rightarrow NH_4^+$

Conjugate pairs

Conjugate base of the acid H_2SO_4

HSO_4-

Conjugate acid of the base HS-

H_2S

Conjugate acid of H_2O

H_3O^+

Conjugate base of H_2O

OH^-

Bronsted-Lowry Acids and Bases

$NH_4^+ + OH^- \leftrightharpoons NH_3 + H_2O$

Acid Base C. Base C. Acid

$C_2H_3O_2^- + H_3O^+ \leftrightharpoons HC_2H_3O_2 + H_2O$

Base Acid C. Acid C. Base

Dissociation Constant (K)

- K_a is the dissociation constant for acid
- K_b is the dissociation constant for base
- The higher the constant (K) the Stronger the acid or base
- Weak acids and bases have a smaller dissociation constant
- **<u>Strong acids produce weak conjugate bases while weak acids produce stronger conjugate bases.</u>**

Monoprotic acids

$HCl \leftrightarrows H^+ + Cl^-$

$HNO_3 \leftrightarrows H^+ + NO_3^-$

$CH_3COOH \leftrightarrows H^+ + CH_3COO^-$

Diprotic acids

$H_2SO_4 \leftrightarrows H^+ + HSO_4^-$

$HSO_4^- \leftrightarrows H^+ + SO_4^{2-}$

Triprotic acids

$H_3PO_4 \leftrightarrows H^+ + H_2PO_4^-$

$H_2PO_4^- \leftrightarrows H^+ + HPO_4^{2-}$

$HPO_4^{2-} \leftrightarrows H^+ + PO_4^{3-}$

**Triprotic acids are generally weak electrolytes and therefore weak acids

Strong Acids and Bases

Strong Acids

- $HClO_4$ (perchloric acid)
- HI (hydroiodic acid)
- HBr (hydrobromic acid)
- HCl (hydrochloric acid)
- H_2SO_4 (sulfuric acid)
- HNO_3 (nitric acid)

Mainly the acid halides!

Strong Bases

- LiOH (lithium hydrox.)
- NaOH (sodium hydrox.)
- KOH (potassium hydrox.)
- RbOH (rubidium hydrox.)
- CsOH (cesium hydrox.)
- $Ca(OH)_2$ (calcium hydrox.)
- $Sr(OH)_2$ (strontium hydrox.)
- $Ba(OH)_2$ (barium hydrox.)

Mainly the bases of alkali metals and some alkaline

Applicable Solution Def's

- **Dissociation:** Process by which the action of a solvent or a change in physical condition, as in pressure or temperature, causes a molecule to split up into simpler groups of atoms, single atoms, or ions
- **Ionization:** Process by which a neutral compound is split into charged particles by action when dissolved in liquid water
- **Equilibrium:** When reactants and products are in a constant ratio. The forward and reverse reactions occur at the same rate when a system is in equilibrium.

Ionization Reactions

- **Completion:** For those that include a **strong** acid or strong base, the reaction will run to completion and can be shown as such with a generic 'yields' symbol (→)

- **Equilibrium:** For those that include a **weak** acid or base or do not go to completion, the reaction can be represented by an equilibrium symbol (⇆)

Equilibrium Constants

- $K_w = [H^+] [OH^-] = 1.0 \times 10^{-14}$
- K_a = [products] / [reactants]

- Weak acids and bases:
  - Easily written, use quadratic (or assumption) to solve by using and I.C.E. chart in examples to follow
- Strong acids and bases: **<u>(cannot be written)</u>**

$$HCl + H_2O \rightarrow Cl^- + H_3O^+$$

$$Ka = [Cl^-][H_3O^+] / [HCl]$$

  - Where $[HCl] \rightarrow 0$, therefore Ka gets very large

Examples of Weak Acids

- HF (hydrofluoric acid)
- H_2CO_3 (carbonic acid)
- H_2S (hydrosulfuric acid)
- H_3BO_3 (boric acid)
- CH_3COOH (acetic acid)
- C_6H_5COOH (benzoic acid)
- $C_6H_4(OH)CO_2COCH_3$ (acetylsalicylic acid)

Acetic acid found in vinegar is a common example of a weak acid found in a household.

Examples of Weak Bases

- NH_3 (ammonia)
- N_2H_4 (hydrazine)
- $(CH_3)_2NH$ (dimethylamine)
- C_5H_5N (pyridine)
- CH_3NH_2 (methylamine)
- $NaHCO_3$ (baking soda)

$NaHCO_3$ otherwise known as baking soda is a common base found in a household.

Helpful Links

- Strong and weak acids
- Acid Base overview

Terms To Know

- pH – the negative of the logarithm to the base ten of the concentration of hydrogen (hydronium) ions in a solution.

$$pH = -\log [\, H^{+}{}_{(aq)} \,]$$

- pOH – the negative of the logarithm to the base ten of the concentration of hydroxide (OH^-) ions in a solution.

$$pOH = -\log [\, OH^{-}{}_{(aq)} \,]$$

Terms To Know (cont'd)

What is an acid?

An acid is a substance from which a hydrogen ion, H^+, can be removed (proton donor)

What is a base?

A base is a substance that is able to remove a hydrogen ion, H^+, from an acid (proton acceptor)

Terms To Know (cont'd)

What makes an acid a weak acid?

A weak acid is an acid that partially ionizes in solution but exists primarily in the form of molecules.

What makes a base a weak base?

A weak base is a base that has a weak attraction for protons.

Terms To Know (cont'd)

- Acid ionization constant (K_a) – equilibrium constant for the ionization of an acid.

- Base ionization constant (K_b) – equilibrium constant for the ionization of a base.

- ion product constant for water (K_w) – equilibrium constant for the ionization of water; $1.0 * 10^{-14}$

Equations to Keep In Mind

- $K_a K_b = K_w$
- $pH = -\log [H^+_{(aq)}]$
- $pOH = -\log [OH^-_{(aq)}]$
- $pH + pOH = 14$ (at SATP)

Steps in Calculating the pH of a solution of a weak acid, given K_a

Step 1 – List the major substances in solution

Step 2 – Write balanced equation for all substances that may produce H^+

Step 3 – Identify the dominant equilibrium, and write the equilibrium constant for the dominant equilibrium.

Step 4 – Complete an ICE table to record changes of the dominant equilibrium.

Step 5 – Substitute the equilibrium concentrations of all the substances of the equation into the acid ionization constant

Step 6 – Assume $[HA]_{initial} - x = [HA]_{initial}$, but only if $[HA]_{initial} / K_a \geq 100$.

Step 7 – solve for x ($[H^+_{(aq)}]$)

Step 8 – Use the 5% rule to check the assumption made in step 6 is correct.

Step 9 – Calculate pH from $[H^+_{(aq)}]$

Sample Problem

Calculate the hydrogen ion concentration and the pH of a 0.10mol/L acetic acid solution. The K_a for acetic acid is $1.8 * 10^{-5.}$

Solution

Step 1 – the major substances in this reaction that may ionize to produce H^+ are H_2O and HC_2H_3O.

Step 2 – $HC_2H_3O_{2\ (aq)} \leftrightarrows H^+_{(aq)} + C_2H_3O_2^-$ $K_a = 1.8 * 10^{-5}$

$H_2O_{(l)} \leftrightarrows H^+_{(aq)} + OH^-$ $K_w = 1.0 * 10^{-14}$

Step 3 – Since the K_a value of $HC_2H_3O_{2\ (aq)}$ is greater than the K_w value of $H_2O_{(l)}$, we can assume that most of the H^+ will be produced by the ionization of $HC_2H_3O_{2\ (aq)}$.

Step 3 still…

Therefore, the ionization of $HC_2H_3O_{2\ (aq)}$ will determine the [H+ (aq)] and the pH of the solution.

For the equilibrium,

$$HC_2H_3O_{2\ (aq)} \leftrightarrows H^+_{\ (aq)} + C_2H_3O_2^{\ -}$$

The equilibrium law equation is

$$K_a = \frac{[\quad]}{\quad}$$

Step 4 – ICE table for ionization of $HC_2H_3O_2\,_{(aq)}$

Step 4	$HC_2H_3O_2\,_{(aq)}$ $\rightleftarrows$	$H^+\,_{(aq)}$	$C_2H_3O_2^{\,-}\,_{(aq)}$
Initial Concentration (mol/L)	0.10	0.00	0.00
Change in Concentration (mol/L)	-x	+x	+x
Equilibrium Concentration (mol/L)	0.10 - x	+x	+x

Step 5

$$\frac{[H^+_{(aq)}]\,[C_2H_3O_2{}^-_{(aq)}]}{[HC_2H_3O_{2\,(aq)}]} = K_a$$

$$\frac{(x)^2}{(0.10 - x)} = 1.8 * 10^{-5}$$

We can simplify the calculation by assuming that, since Ka is so small, $HC_2H_3O_{2\,(aq)}$ will ionize very little and the value of x is expected to be very small. But to make sure that the assumption is valid, the Hundred Rule must be used first to validate the assumption

Step 6

$$\frac{[HA]initial}{K_a} = \frac{0.10}{1.8 * 10^{-5}} = 5.6 * 10^{3}$$

Since $5.6 * 10^3 > 100$, we can assume that $0.10 - x = 0.10$.

Step 7

The equilibrium equation becomes

$$\frac{x^2}{0.10} = 1.8 * 10^{-5}$$

$$x^2 = 1.8 * 10^{-6}$$

$$x = 1.3 * 10^{-3}$$

Step 8

We must now validate the approximation of 0.10 – x = 0.10. Since Ka values are usually known to an accuracy of $\pm$ 5%, we will use this figure to validate the assumption.

$$\frac{x}{[HA]initial} * 100\% \leq 5\%$$

$$\frac{1.3 * 10^{-3}\ mol/L}{0.10\ mol/L} * 100\% \leq 5\%$$

$$1.3\% \leq 5\%$$

Therefore the assumption can be considered valid and the value of the calculated value of x is acceptable.

Step 9

$[H^+_{(aq)}] = x$

$[H^+_{(aq)}] = 1.3 * 10\text{-}3$ mol/L

$pH = -\log [H^+_{(aq)}]$

$= -\log (1.3 * 10\text{-}3 \text{ mol/L})$

$= 2.89$

Therefore the pH of a 0.10 mol/L acetic acid solution is 2.89.

Steps in Calculating the pH of a solution of a weak base, given K_b

Step 1 – List the major substances in solution

Step 2 – Write balanced equation for all substances that may produce OH^+

Step 3 – Identify the dominant equilibrium, and write the equilibrium constant for the dominant equilibrium.

Step 4 – Complete an ICE table to record changes of the dominant equilibrium.

Step 5 – Substitute the equilibrium concentrations of all the substances of the equation into the acid ionization constant

Step 6 – Assume $[B]_{initial} - x = [B]_{initial}$, but only if $[B]_{initial} / K_b \geq 100$.

Step 7 – solve for x ($[\ OH^-_{(aq)}\ ]$)

Step 8 – Use the 5% rule to check the assumption made in step 6 is correct.

Step 9 – Calculate pOH from $[\ OH^+_{(aq)}\ ]$

Step 10 – Calculate the pH by substituting the value of pOH into pH + pOH = 14

Sample Problem

Calculate the pH of a 0.100mol/L aqueous solution of hydrazine, $N_2H_{4\,(aq)}$, a weak base given that K_b for hydrazine is $1.7 * 10^{-6}$

Solution

Step 1 – the major substances in solution are $N_2H_{4\,(aq)}$ and $H_2O_{\,(aq)}$.

Step 2

$$N_2H_{4\,(aq)} + H_2O_{\,(l)} \rightleftharpoons N_2H_5^{+}{}_{(aq)} + OH^{-}{}_{(aq)} \qquad K_b = 1.7 * 10^{-6}$$

$$H_2O_{\,(l)} \rightleftharpoons H^{+}{}_{(aq)} + OH^{-}{}_{(aq)} \qquad K_w = 1.0 * 10^{-14}$$

Step 3 – Since the Kw << Kb, the N_2H_4 equilibrium will predominate in the solution, and it will be assumed that all OH^- ions present in the solution will be produced by the $N_2H_{4\,(aq)}$ equilibrium.

For the equilibrium,

$$HC_2H_3O_{2\,(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O_2^{\,-}$$

The equilibrium law equation is

$$\frac{[N_2H_5^+{}_{(aq)}]\,[OH^-{}_{(aq)}]}{[N_2H_{4(aq)}]} = K_b$$

Step 4 – ICE table for ionization of $N_2H_{4\,(aq)}$

	$N_2H_{4\,(aq)}$ +	$H_2O_{(l)}$ $\rightleftharpoons$	$N_2H_5^+{}_{(aq)}$	$OH^-{}_{(aq)}$
Initial Concentration (mol/L)	0.100	-	0.00	0.00
Change in Concentration (mol/L)	-x	-	+x	+x
Equilibrium Concentration (mol/L)	0.100 - x	-	+x	+x

Step 5

$$\frac{[N_2H_5^+{}_{(aq)}]\,[OH^-{}_{(aq)}]}{[N_2H_{4\,(aq)}]} = K_b$$

$$\frac{(x)^2}{(0.10 - x)} = 1.7 * 10^{-6}$$

The Hundred Rule will then be used to see whether a simplifying assumption may be made.

Step 6

$$\frac{[HA]initial}{K_a} = \frac{0.10}{1.7 * 10^{-6}}$$

$$= 5.9 * 10^{-4}$$

Since $5.9 * 10^{-4} >> 100$, it is safe to assume that 0.100 – x = 0.1

Step 7

The equilibrium equation becomes:

$$\frac{x^2}{0.10} = 1.7 * 10^{-6}$$

$$x^2 = 1.7 * 10^{-7}$$

$$x = 4.12 * 10^{-4}$$

Step 8

$$\frac{x}{[HA]initial} * 100\% \leq 5\%$$

$$\frac{4.12 * 10^{-4} \text{ mol/L}}{0.10 \text{ mol/L}} * 100\% \leq 0.41\%$$

Since 0.41% << 5%, the assumption made is justified.

Step 9

$x = 4.12 * 10^{-4}$

$[OH^-_{(aq)}] = 4.12 * 10^{-4}$

$pOH = -\log [OH^-_{(aq)}]$

$= -\log [4.12 * 10^{-4}]$

$= 3.38$

Step 10

pH + pOH = 14

pH = 14.00 - pOH

pH = 14.00 – 3.38

pH = 10.62

Therefore the pH of a 0.100 mol/L hydrazine is 10.62.

Practice Questions

1) Determine the pH of a solution with a hydrogen concentration of $4.8 * 10^{-11}$?

2) Determine the pH of a solution with a hydroxide concentration of $3.0 * 10^{-6}$?

3) Determine the pH of a 0.01 mol/L solution of chloracetic acid given $Ka = 1.36 * 10^{-3}$?

4) Determine the pH of a solution of morphine that has a concentration of 0.01mol/L, given that Kb for morphine is $7.5 * 10^{-7}$?

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